**AP Chemistry**

**Acids and Bases (Chapters 16 and 17) Practice Problems**

**Multiple Choice Questions**

*Identify the letter of the choice that best completes the statement or answers the question. You may not use a calculator to answer these questions, but you may refer to your periodic table and formula sheets.*

\_\_\_\_ 1. How can 100. mL of sodium hydroxide solution with a pH of 13.00 be converted to a sodium hydroxide solution with a pH of 12.00?

|  |  |
| --- | --- |
| a. | By diluting the solution with distilled water to a total volume of 108 mL |
| b. | By diluting the solution with distilled water to a total volume of 200 mL |
| c. | By diluting the solution with distilled water to a total volume of 1.00 L |
| d. | By adding 100. mL of 0.10 *M* HCl |

\_\_\_\_ 2. Ascorbic acid, H2C6H6O6*(s)*, is a diprotic acid with K1 = 7.910-5 and K2 = 1.610-12. In a 0.005 *M* aqueous solution of ascorbic acid, which of the following species is present in the lowest concentration?

|  |  |  |  |
| --- | --- | --- | --- |
| a. | H3O+*(aq)* | c. | H2C6H6O6*(aq)* |
| b. | C6H6O62-*(aq)* | d. | HC6H6O6-*(aq)* |

\_\_\_\_ 3. A pure, white crystalline solid dissolves in water to yield a basic solution that liberates a gas when excess acid is added to it. On the basis of this information, the solid could be

|  |  |  |  |
| --- | --- | --- | --- |
| a. | KNO3 | c. | KOH |
| b. | K2CO3 | d. | KHSO4 |

\_\_\_\_ 4. The pH of a solution prepared by the addition of 10. mL of 0.002M KOH(aq) to 10. mL of distilled water is closest to

|  |  |  |  |
| --- | --- | --- | --- |
| a. | 12 | d. | 4 |
| b. | 11 | e. | 3 |
| c. | 10 |

\_\_\_\_ 5. What is the H+*(aq)* concentration in 0.05 *M* HCN*(aq)* ? (The *Ka* for HCN is 5.0x10–10.)

|  |  |  |  |
| --- | --- | --- | --- |
| a. | 2.5x10–11 *M* | c. | 5.0x10–10 *M* |
| b. | 2.5x10–10 *M* | d. | 5.0x10–6 *M* |

\_\_\_\_ 6.

HC2H3O2*(aq)* + CN–*(aq)* ⇌ HCN*(aq)* + C2H3O2–*(aq)*

The reaction represented above has an equilibrium constant equal to 3.7x104. Which of the following can be concluded from this information?

|  |  |
| --- | --- |
| a. | CN–*(aq)* is a stronger base than C2H3O2–*(aq)*. |
| b. | HCN*(aq)* is a stronger acid than HC2H3O2 *(aq)* . |
| c. | The conjugate base of CN–*(aq)* is C2H3O2–*(aq)*. |
| d. | The equilibrium constant will increase with an increase in temperature. |

\_\_\_\_ 7.

HSO4– + H2O ⇌ H3O+ + SO42–

In the equilibrium represented above, the species that act as bases include which of the following?

I. HSO4– II. H2O III. SO42–

|  |  |  |  |
| --- | --- | --- | --- |
| a. | II only | c. | II and III |
| b. | III only | d. | I and III |

\_\_\_\_ 8. All of the following species can function as Bronsted–Lowry bases in solution EXCEPT.

|  |  |  |  |
| --- | --- | --- | --- |
| a. | H2O | c. | S2– |
| b. | NH3 | d. | NH4+ |

\_\_\_\_ 9. A 0.20–molar solution of a weak monoprotic acid, HA, has a pH of 3.00. The ionization constant of this acid is

|  |  |  |  |
| --- | --- | --- | --- |
| a. | 5.0x10–7 | c. | 5.0x10–6 |
| b. | 2.0x10–7 | d. | 5.0x10–3 |

\_\_\_\_ 10.

If the acid dissociation constant, Ka, for an acid HA is 8x10–4 at 25oC, what percent of the acid is dissociated in a 0.50–molar solution of HA at 25oC?

|  |  |  |  |
| --- | --- | --- | --- |
| a. | 0.08% | c. | 1% |
| b. | 4% | d. | 2% |

\_\_\_\_ 11. Which of the following mixtures would be considered a buffer?

|  |  |  |  |
| --- | --- | --- | --- |
| a. | 0.10 *M* HCl + 0.10 *M* NaCl | c. | 0.10 *M* HBr + 0.10 *M* NaBr |
| b. | 0.10 *M* HF + 0.10 *M* NaF | d. | 0.10 *M* HClO2 + 0.10 *M* NaCl |

\_\_\_\_ 12. The net ionic equation for the reaction that occurs during the titration of nitrous acid with sodium hydroxide is

|  |  |
| --- | --- |
| a. | HNO2 + Na+ + OH–  NaNO2 + H2O |
| b. | HNO2 + NaOH  Na+ + NO2– + H2O |
| c. | H+ + OH– H2O |
| d. | HNO2 + OH–  NO2– + H2O |

**Matching**

A lettered choice may be used once, more than once, or not at all.

Questions 13-15 refer to the following table of ionization constants

Ionization constants:

CH3COOH = 1.8x10–5

NH3 = 1.8x10–5

H2CO3: K1 = 4x10–7

K2 = 4x10–11

|  |  |
| --- | --- |
| a. | a solution with a pH less than 7 that is not a buffer solution |
| b. | a buffer solution with a pH between 4 and 7 |
| c. | a buffer solution with a pH between 7 and 10 |
| d. | a solution with a pH greater than 7 that is not a buffer solution |
| e. | a solution with a pH of 7 |

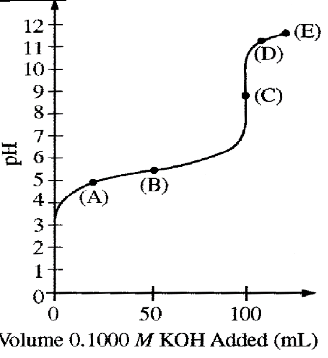
\_\_\_\_ 13. A solution prepared to be initially 1 M in NaCl and 1 M in HCl.

\_\_\_\_ 14. A solution prepared to be initially 1 M in Na2CO3 and 1 M in CH3COONa

\_\_\_\_ 15. A solution prepared to be initially 0.5 M in CH3COOH and 1 M in CH3COONa

Questions 16-19 refer to the information provided below.

A solution of a weak monoprotic acid is titrated with a solution of a strong base, KOH. Consider the points labeled A through E on the titration curve that results, as shown below.



|  |  |  |  |
| --- | --- | --- | --- |
| a. | A | d. | D |
| b. | B | e. | E |
| c. | C |

\_\_\_\_ 16. The point at which the moles of the added strong base are equal to the moles of the weak acid initially present

\_\_\_\_ 17. The point at which the pH is closest to that of the strong base being added

\_\_\_\_ 18. The point at which the concentrations of the weak acid and its conjugate base are approximately equal

\_\_\_\_ 19. The point at which the pH of the solution equals the pKa of the weak acid

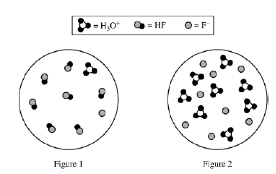
**Free Response Questions**

20. The ionization of HF(aq) in water is represented by the equation below. (2018)

HF*(aq)* + H2O*(l)* ⇌ F-*(aq)* + H3O+*(aq)*

In a 0.0350 M HF(aq) solution, the percent ionization of HF is 13.0 percent.

a) Two particulate representations of the ionization of HF molecules in the 0.0350 M HF(aq) solution are shown below in Figure 1 and Figure 2. Water molecules are not shown. Explain why the representation of the ionization of HF molecules in water in Figure 1 is more accurate than the representation of in Figure 2. (The key below identifies the particles in the representations.) (1 pt)



b) Use the percent ionization data above to calculate the value of Ka of HF. (2 pts)

c) If 50.0 mL of distilled water is added to 50.0 mL of 0.035M HF(aq), will the percent ionization of HF(aq) increase, decrease or remain the same: Justify your answer with an explanation or calculation.

21. **2003**

C6H5NH2*(aq)* + H2O*(l)* ⇌ C6H5NH3+*(aq)* + OH-*(aq)*

Aniline, a weak base, reacts with water according to the reaction represented above.

(a) Write the equilibrium constant expression, K*b*, for the reaction represented above.

(b) A sample of aniline is dissolved in water to produce 25.0 mL of 0.10 *M* solution. The pH of the solution is 8.82. Calculate the equilibrium constant, K*b*, for this reaction.

(c) The solution prepared in part (b) is titrated with 0.10 *M* HCl. Calculate the pH of the solution when 5.0 mL of the acid has been added.

(d) Calculate the pH at the equivalence point of the titration in part (c).

(e) The p*K*a values for several indicators are given below. Which of the indicators listed is most suitable for this titration? Justify your answer.

|  |  |
| --- | --- |
| Indicator | p*Ka* |
| Erythrosine | 3 |
| Litmus | 7 |
| Thymolphthalein | 10 |

22. Phenol is a weak acid that partially dissociates in water, according to the equation below. (2016)

C6H5OH(aq) + H2O() ⇌ C6H5O-(aq) + H3O+(aq) Ka = 1.12 x 10-10

a) What is the pH of a 0.75M C6H5OH(aq) solution? (2 pts)

b) For a certain reaction involving C6H5OH(aq) to proceed at a particular rate, the phenol must be primarily in its deprotonated form, C6H5O-(aq). In order to ensure that the C6H5OH(aq) is deprotonated, the reaction must be conducted in a buffered solution. On the number scale below, circle each pH for which more than 50% of the phenol molecules are in the deprotonated form (C6H5O-(aq)). Justify your answer. (2 pts)

1 2 3 4 5 6 7 8 9 10 11 12 13 14

23. **2002 A Required**

HOBr*(aq)* ⇌ H+*(aq)* + OBr–*(aq)* *Ka* = 2.3  10–9

Hypobromous acid, HOBr, is a weak acid that dissociates in water, as represented by the equation above.

(a) Calculate the value of [H+] in an HOBr solution that has a pH of 4.95.

(b) Write the equilibrium constant expression for the ionization of HOBr in water, then calculate the concentration of HOBr*(aq)* in an HOBr solution that has [H+] equal to 1.8  10-5 *M*.

(c) A solution of Ba(OH)2 is titrated into a solution of HOBr.

(i) Calculate the volume of 0.115 *M* Ba(OH)2*(aq)* needed to reach the equivalence point when titrated into a 65.0 mL sample of 0.146 *M* HOBr*(aq)*.

(ii) Indicate whether the pH at the equivalence point is less than 7, equal to 7, or greater than 7. Explain.

(d) Calculate the number of moles of NaOBr*(s)* that would have to be added to 125 mL of 0.160 *M* HOBr to produce a buffer solution with [H+] = 5.00  10-9 *M*. Assume that volume change is negligible.

(e) HOBr is a weaker acid than HBrO3. Account for this fact in terms of molecular structure.

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24. 1996 A

HOCl ⇌ OCl- + H+

Hypochlorous acid, HOCl, is a weak acid commonly used as a bleaching agent. The acid-dissociation constant, *Ka,* for the reaction represented above is 3.2x10-8.

(a) Calculate the [H+] of a 0.14-molar solution of HOCl.

(b) Write the correctly balanced net ionic equation for the reaction that occurs when NaOCl is dissolved in water and calculate the numerical value of the equilibrium constant for the reaction.

(c) Calculate the pH of a solution made by combining 40.0 milliliters of 0.14-molar HOCl and 10.0 milliliters of 0.56-molar NaOH.

(d) How many moles of solid NaOH must be added to 50.0 mL of 0.20-molar HOCl to obtain a buffer solution that has a pH of 7.49? Assume that the addition of the solid NaOH results in a negligible change in volume.

(e) Household bleach is made by dissolving chlorine gas in water, as represented below.

Cl2*(g)* + H2O → H+ + Cl- + HOCl*(aq)*

Calculate the pH of such a solution if the concentration of HOCl in the solution is 0.065 molar.

**Acids and Bases (Chapters 16 and 17) Practice Problems**

**Answer Section**

**MULTIPLE CHOICE**

1. ANS: C PTS: 1

2. ANS: B PTS: 1

3. ANS: B PTS: 1

4. ANS: B PTS: 1

5. ANS: D PTS: 1

6. ANS: A PTS: 1

7. ANS: C PTS: 1

8. ANS: D PTS: 1

9. ANS: C PTS: 1

10. ANS: B PTS: 1

11. ANS: B PTS: 1

12. ANS: D PTS: 1

**MATCHING**

13. ANS: A PTS: 1

14. ANS: D PTS: 1

15. ANS: B PTS: 1

16. ANS: C PTS: 1

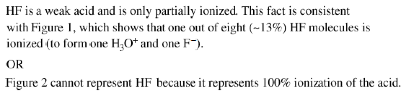
17. ANS: E PTS: 1

18. ANS: B PTS: 1

19. ANS: B PTS: 1

**PROBLEM**

20. ANS:

a) 

b) 6.81 x 10-4 for original problem

ICE problem--you cannot assume x is neglible 1 pt for [H3O+], 1 pt for Ka value

c) % ionization will increase Q = 1/2 K so system will proceed right, or use new concentrations to recalculate [H+] using Ka expression

PTS: 1

21. ANS:

Answer:

(a) *K*b = 

(b) pOH = 14 – pH = 14 – 8.82 = 5.18

-log[OH–] = 5.18; [OH–] = 6.6110–6 *M*

[OH–] = [C6H5NH3+] (can set up an ice table here, may assume that x << 0.10 but not required)

*K*b =  = 4.410–10

(c) (0.025 L)(0.10 mol/L) = 0.025 moles weak base initially in solution

moles H+ added from the HCl: (0.0050 L)(0.10 mol/L) = 0.0050 moles H+ added

LR problem! H+ reacts with weak base in a 1:1 ratio, so 0.020 moles of weak base remain and we make 0.0050 moles of the conjugate acid (so we actually have a buffer system). The total solution volume is 30.0 mL at this point



[H+] =  = 5.610–6; pH = 5.26

(d) when neutralized, there are 0.025 mol of C6H5NH3+ in 50.0 mL of solution, giving a [C6H5NH3+] = 0.050 *M*

this cation will partially ionize according to the following equilibrium:

C6H5NH3+*(aq)*  C6H5NH2*(aq)* + H+*(aq)*

*Note that this is the conjugate acid of the weak base we started with....so we need to calculate Ka from the Kb calculated in part (b)*



at equilibrium, [C6H5NH2] = [H+] = *X*

[C6H5NH3+] = (0.050–*X*)

** = *K*a = 2.310-5

*X* = 1.0610–3 = [H+]

pH = –log[H+] = 2.98

(e) erythrosine; the indicator will change color when the pH is near its p*K*a, since the equivalence point is near pH 3, the indicator must have a p*K*a near this value.

PTS: 1

22. ANS:

A) ice table, Ka expression

x = 9.2 x 10-6 M (1 pt for correct setup, calculation)

pH = 5.04 (1 pt based on correct setup & calculation)

B) Numbers 10-14 should be circled (1 pt)

When pH >pKa, the deprotonated form (i.e. conjugate base) form will predominate

pKa = 9.95 therefore at pH 10 and above, [C6H5O-]>[C6H5OH]

PTS: 1

23. ANS:

(a) pH = -log[H+]; 4.95 = -log[H+]

[H+] = 1.12  10-5 M

(b) *K*a = = 2.3  10-9

[H+] = [OBr–] = 1.8  10-5 *M*

= 2.3  10-9

X = [HOBr] = 0.14 *M*

(c) (i) (0.650L HOBr)(0.146 mol/L) (1 mol Ba(OH)2/2 mol HOBr) = 9.49 x 10-3 mol Ba(OH)2

( 9.49 x 10-3 mol)(1L/0.115 mol) - 41.3 mL

41.3 mL Note the 2:1 mole ratio--write a balanced equation!

(ii) pH > 7; the dominant species at the equivalence point is the conjugate base, which will increase the pH

(d) [OBr-] = 

(e) very electronegative oxygen is able to draw electrons away from the bromine and weaken the O–H bond, making it easier for the hydrogen ion “to leave”.



PTS: 1

24. ANS:

(a) *Ka* =  = 3.2x10-8

*X* = amount of acid that ionizes = [OCl-] = [H+]

(0.14 - X) = [HOCl] that remains unionized

3.2x10-8 =  ; *X* = 6.7x10-5 M = [H+]

(b) NaOCl*(s)* + H2O --> Na+*(aq)* + HOCl*(aq)* + OH-*(aq)*

*Kb* = = 3.1x10-7

(c) [ ]o after dilution but prior to reaction:

[HOCl] = 0.14 *M* x  = 0.11 *M*

[OH-] = 0.56 *M* x  = 0.11 *M*

Equivalence point reached. [OH-] ~ [HOCl]

*Kb* =  = 3.1x10-7

[OH-] = 1.8x10-4 ; pOH = 3.7

pH = 14 - 3.7 = 10.3

(d) at pH 7.49, the [H+] = 10-7.49 = 3.24x10-8 *M*

when the solution is half-neutralized, pH = p*Ka* and  = 1

 x 50.0 mL = 10.0 mmol HOCl

half this amount, or 5.0 mmol of NaOH added.

(e) 1 mol H+ for every 1 mole of HOCl produced

[H+] ~ [HOCl] = 0.065 *M*

pH = - log (0.065) = 1.2

PTS: 1